13.1 Structure of the Nucleus

SECTION EXPECTATIONS

- Define and describe the concepts and units related to the present-day understanding of the nature of elementary particles (e.g., mass-energy equivalence).
- Apply quantitatively the laws of conservation of mass and energy, using Einstein's mass-energy equivalence.

KEY TERMS

- proton
- neutron
- nucleon
- chemical symbol
- atomic number
- atomic mass number
- nucleon number
- strong nuclear force
- nuclide
- isotope
- mass defect
- atomic mass unit

Excitement was high in the scientific community in the early 1900s when Ernest Rutherford (1871–1937) proposed his model of the nucleus and Niels Bohr (1885–1962) developed a model of the atom that explained the spectrum of hydrogen. As is the case with many scientific breakthroughs, however, the answers to a few questions gave rise to many more.

The most obvious question arose from the realization that a great amount of positive charge was concentrated in a very small space inside the nucleus. The strength of the Coulomb repulsive force between like charges had long been established. Two positive charges located as close together as they would have to be in Rutherford's model of the nucleus would exert a mutual repulsive force of about 50 N on each other. For such tiny particles, this is a tremendous force. There had to be another, as yet unidentified, attractive force that was strong enough to overcome the repulsive Coulomb force. What is the nature of the particles that make up the nucleus and what force holds them together?

Protons and Neutrons

Again, it was Rutherford who discovered — and eventually named — the proton. When he was bombarding nitrogen gas with alpha particles, Rutherford detected the emission of positively charged particles with the same properties as the hydrogen nucleus. As evidence accumulated, it became apparent to physicists that the **proton** was identical to the hydrogen nucleus and was the fundamental particle that carried a positive charge, equal in magnitude to the charge on the electron and with a mass 1836 times as great as the mass of an electron. The positive charge of all nuclei consisted of enough protons to account for the charge.

Using the principle on which the mass spectrometer is based (refer to Chapter 8, Fields and Their Applications), several physicists discovered that the mass of most nuclei was roughly twice the size of the number of protons that would account for the charge. Rutherford encouraged the young physicists in his laboratory to search for a neutrally charged particle that could account for the excess mass of the nucleus. Finally, in 1932, English physicist James Chadwick (1891–1974) discovered such a particle. That particle, now called the **neutron**, has a mass that is nearly the same as that of a proton. The proton, neutron, and electron now account for all of the mass and charge of the atom. Since protons and neutrons have many characteristics in common, other than the charge, physicists call them **nucleons**.

Table 13.1 Properties of Particles in the Atom

Particle	Mass (kg)	Charge (C)
proton	$1.672\ 614 \times 10^{-27}$ kg	$+1.602 \times 10^{-19}$ C
neutron	$1.674\ 920 \times 10^{-27}$ kg	0 C
electron	$9.109\ 56 \times 10^{-31}$ kg	-1.602×10^{-19} C

Representing the Atom

As physicists and chemists learned more about the nucleus and atoms, they needed a way to symbolically describe them. The following symbol convention communicates much information about the particles in the atom.

$${}^{A}_{Z}X$$

X is the **chemical symbol** for the element to which the atom belongs. For example, the symbol for carbon is C, while the symbol for krypton is Kr.

Z is the **atomic number**, which represents the number of protons in the nucleus and is also the charge of the nucleus.

A is the **atomic mass number**, the total number of protons and neutrons in the nucleus. Since the particles in the nucleus are called "nucleons," the atomic mass number is sometimes called the **nucleon number**.

If N represents the number of neutrons in a nucleus, then

$$A = Z + N$$

The atomic number (Z) also indicates the number of electrons in the neutral atom, since one electron must be in orbit outside the nucleus for each proton inside the nucleus. In addition, the manner in which atoms chemically interact with each other depends on the arrangement of their outer electrons. This is influenced in turn by the atomic number. Since all atoms of an element behave the same chemically, all atoms of a given element must have the same atomic number. For example, all carbon atoms have an atomic number of 6 and all uranium atoms have an atomic number of 92.

The Strong Nuclear Force

By the end of the 1930s, physicists were beginning to accumulate data about the elusive force that holds the nucleus together. They discovered that any two protons (p↔p), two neutrons (n↔n), or a proton and a neutron (p↔n) attract each other with the most potent force known to physicists — the **strong nuclear force**.

When two protons are about 2 fm (femtometres: 2×10^{-15} m) apart, the nuclear force is roughly 100 times stronger than the repulsive Coulomb force. However, at 3 fm of separation, the nuclear force is almost non-existent. Whereas the gravitational and electrostatic forces have an unlimited range — both follow a $1/r^2$ law — the nuclear force has an exceptionally short range, which is roughly the diameter of a nucleon. Therefore, inside the nucleus, the nuclear force acts only between adjacent nucleons. When the separation distance between nucleons decreases to about 0.5 fm, the nuclear force becomes repulsive. This repulsion possibly occurs because nucleons cannot overlap. Estimates of the radius of a nucleon range from 0.3 fm to 1 fm. Figure 13.1 shows a graph of an approximated net force between two protons, relative to their separation distance.

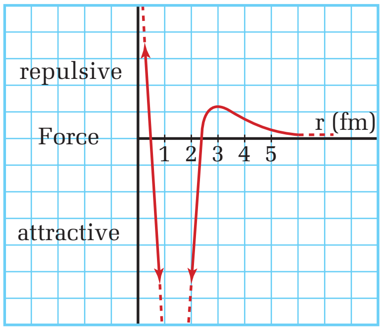

Figure 13.1 Assume that one proton is at the origin of the co-ordinate system and another proton approaches it. As the second proton approaches, it experiences a repulsive Coulomb force. If the second proton has enough energy to overcome the repulsion, it will reach a point where the net force is zero. Beyond that point, the strong nuclear force attracts it strongly.

Stability of the Nucleus

If the nuclear force is so strong, why cannot nucleons come together to form nuclei of ever-increasing size? The short range of the nuclear force accounts for this. When a nucleus contains more than approximately 20 nucleons, the nucleons on one side of the nucleus are so far from those on the opposite side that they no longer attract each other. However, the repulsive Coulomb force between protons is still very strong. Figure 13.2 shows the number of protons (Z) and neutrons (N) in all stable nuclei.

As you can see in Figure 13.2, the number of neutrons and protons is approximately equal up to a total of 40 nucleons. For example, oxygen has 8 protons and 8 neutrons and calcium has 20 protons and 20 neutrons $\left({}^{40}_{20}\text{Ca}\right)$. Beyond 40 nucleons, the ratio of neutrons to protons increases gradually up to the largest element. One form of uranium has 92 protons and 146 neutrons $\left({}^{238}_{92}\text{U}\right)$. This unbalanced ratio results in more nucleons experiencing the attractive force for each pair of protons experiencing the repulsive Coulomb forces. This combination appears to stabilize larger nuclei. Nuclei that do not lie in the range of stability will disintegrate.

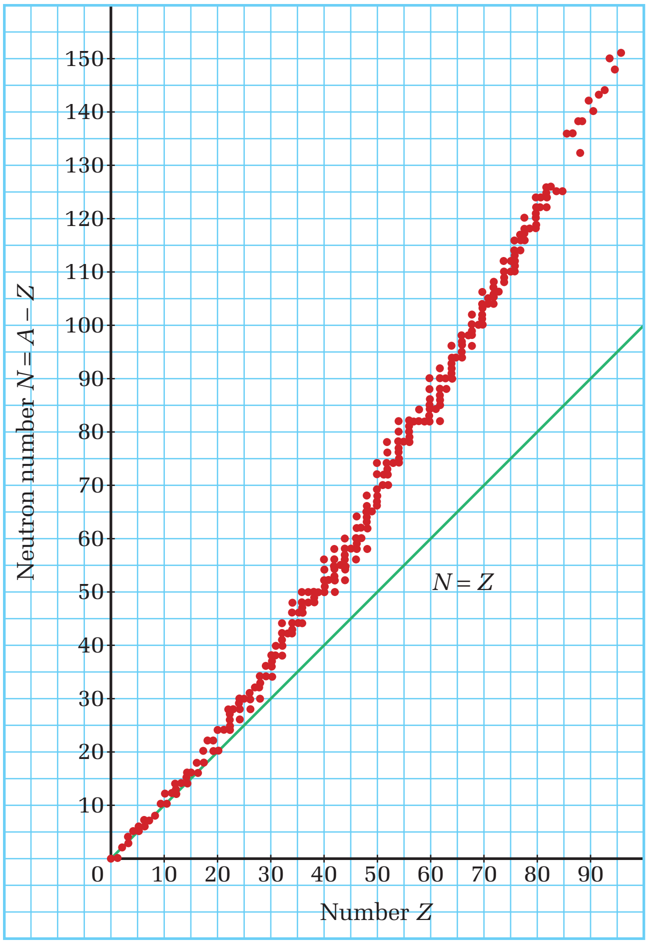

Figure 13.2 Each dot represents a stable nucleus, with the number of neutrons shown on the vertical axis and the number of protons on the horizontal axis.

Nuclides and Isotopes

Each dot in Figure 13.2 represents a unique stable nucleus with a different combination of protons and neutrons. Physicists call these unique combinations **nuclides**. The many columns of vertical dots indicate that several nuclides have the same number of protons. Since the number of protons determines the identity of the element, all of the nuclides in a vertical column are different forms of the same element, differing only in the number of neutrons. These sets of nuclides are called **isotopes**. For example, nitrogen, with 7 protons, might have 7 neutrons $\left({}^{14}_{7}\mathrm{N}\right)$ or 8 neutrons $\left({}^{15}_{7}\mathrm{N}\right)$. Figure 13.3 illustrates isotopes of hydrogen and helium.

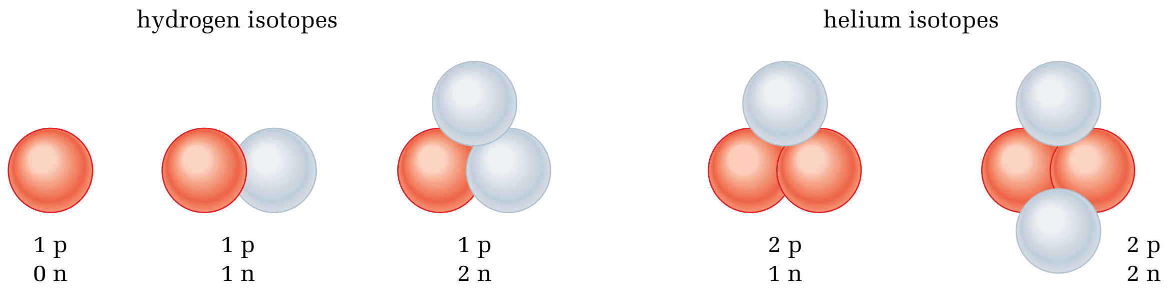

Figure 13.3 The isotopes of hydrogen are the only isotopes to which physicists have given different names: ${}^{2}_{1}H$ is called "deuterium" and ${}^{3}_{1}H$ is called "tritium." For most isotopes, physicists simply use the atomic mass number to describe the isotope: ${}^{3}_{2}He$ is called "helium-3" and ${}^{4}_{2}He$ is called "helium-4."

Nuclear Binding Energy and Mass Defect

When you consider the strength of the nuclear force, you realize that it would take a tremendous amount of energy to remove a nucleon from a nucleus. For the sake of comparison, recall that it takes 13.6 eV to ionize a hydrogen atom, which is the removal of the electron. To remove a neutron from ${}^{4}_{2}He$ would require more than 20 million eV (20 MeV). The amount of energy required to separate all of the nucleons in a nucleus is called the binding energy of the nucleus. Figure 13.4 shows the average binding energy per nucleon plotted against atomic mass number A.

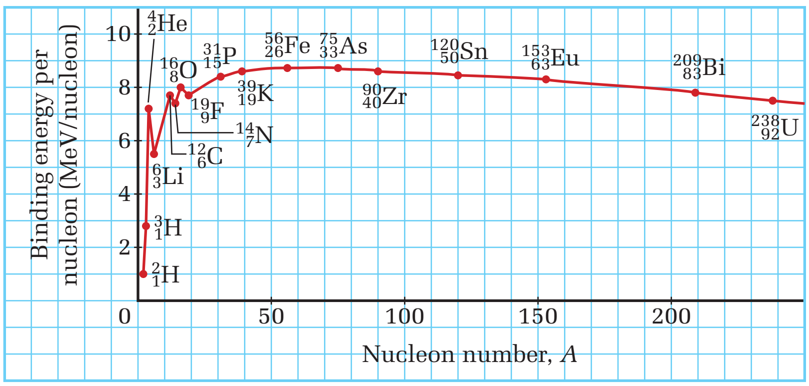

Figure 13.4 The average binding energy per nucleon is calculated by determining the total binding energy of the nucleus and dividing by the number of nuclides.

WEB LINK

www.mcgrawhill.ca/links/physics12

You can find many properties, including the mass, of all stable nuclides and many unstable nuclides in charts of the nuclides on the Internet. Just go to the above Internet site and click on **Web Links**.

Imagine that you were able to remove a neutron from ${}^{4}_{2}He$. What would happen to the 20 MeV of energy that you had to add in order to remove the neutron? The answer lies in Einstein's special theory of relativity: Energy is equivalent to mass. If you look up the masses of the nuclides, you would find that the mass of ${}^{4}_{2}He$ is

smaller than the sum of the masses of ${}^{3}_{2}He$ plus a neutron $\left({}^{1}_{0}n\right)$. The energy that was added to remove a neutron from ${}^{4}_{2}He$ became mass. This difference between the mass of a nuclide and the sum of the masses of its constituents is called the **mass defect**. Einstein's equation $E = \Delta mc^2$ allows you to calculate the energy equivalent of the mass defect, Δm.

When dealing with reactions involving atoms or nuclei, expressing masses in kilograms can be cumbersome. Consequently, physicists defined a new unit — the **atomic mass unit** (u). One atomic mass unit is defined as $\frac{1}{12}$ the mass of the most common isotope of carbon $\left({}^{12}_{6}C\right)$. This gives a value for the atomic mass unit of $1\ u = 1.6605 \times 10^{-27}$ kg. Table 13.2 lists the masses of the particles in atoms.

Table 13.2 Masses of Common Elementary Particles

Particle	Mass (kg)	Mass (u)
electron	$9.109\ 56 \times 10^{-31}$ kg	0.000 549 u
proton	$1.672\ 614 \times 10^{-27}$ kg	1.007 276 u
neutron	$1.674\ 920 \times 10^{-27}$ kg	1.008 665 u

SAMPLE PROBLEM

Calculate the Binding Energy of a Nucleus

Determine the binding energy in electron volts and joules for an iron nucleus of $\left({}^{56}_{26}Fe\right)$, given that the nuclear mass is 55.9206 u.

Conceptualize the Problem

- The *energy equivalent* of the *mass defect* is the *binding energy* for the nucleus.
- The *mass defect* is the *difference* of the mass of the *nucleus* and the *sum* of the masses of the *individual particles*.

Identify the Goal

The binding energy, E, of ${}^{56}_{26}Fe$

Identify the Variables and Constants

Known	Implied	Unknown
$m_{nucleus} = 55.9206$ u	$c = 2.998 \times 10^{8}\ \frac{m}{s}$	N
$A = 56$	$m_p = 1.007\ 276$ u	Δm
$Z = 26$	$m_n = 1.008\ 665$ u	E

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Develop a Strategy

Calculate the number of neutrons.

$$N = A - Z$$
$$N = 56 - 26$$
$$N = 30$$

Determine the total mass of the separate nucleons by finding the masses of the protons and neutrons and adding them together.

$$m_{p(total)} = (26)(1.007\ 276\ \text{u})$$
$$m_{p(total)} = 26.189\ 176\ \text{u}$$

Find the mass defect by subtracting the mass of the nucleus from the total nucleon mass.

$$m_{n(total)} = (30)(1.008\ 665\ \text{u})$$
$$m_{n(total)} = 30.259\ 95\ \text{u}$$
$$m_{total} = 26.189\ 176\ \text{u} + 30.259\ 95\ \text{u}$$
$$m_{total} = 56.449\ 126\ \text{u}$$
$$\Delta m = 56.449\ 126\ \text{u} - 55.9206\ \text{u}$$
$$\Delta m = 0.528\ 526\ \text{u}$$

Convert this mass into kilograms.

$$\Delta m = (0.528\ 526\ \cancel{\text{u}})\left(1.6605 \times 10^{-27}\ \frac{\text{kg}}{\cancel{\text{u}}}\right)$$
$$\Delta m = 8.7762 \times 10^{-28}\ \text{kg}$$

Find the energy equivalent of the mass defect.

$$E = \Delta mc^2$$

Find the energy in electron volts.

$$\Delta E = \left(8.7762 \times 10^{-28}\ \text{kg}\right)\left(2.998 \times 10^{8}\ \frac{\text{m}}{\text{s}}\right)^2$$
$$\Delta E = 7.888 \times 10^{-11}\ \text{J}$$
$$\Delta E = \frac{7.888 \times 10^{-11}\ \cancel{\text{J}}}{1.602 \times 10^{-19}\ \frac{\cancel{\text{J}}}{\text{eV}}}$$
$$\Delta E = 4.9239 \times 10^{8}\ \text{eV}$$

The binding energy of the nucleus is 4.924×10^8 eV, or 7.888×10^{-11} J.

Validate the Solution

The binding energy of a nucleus should be extremely small.
You would expect the binding energy per nucleon to be about 8 MeV.

$$\frac{4.93 \times 10^8\ \text{eV}}{56} = 8.79 \times 10^6\ \text{eV} = 8.79\ \text{MeV}$$

PRACTICE PROBLEMS

1. Determine the mass defect for ${}^{8}_{4}\text{Be}$ with a nuclear mass of 8.003 104 u.

2. Determine the binding energy for ${}^{3}_{2}\text{He}$ with a nuclear mass of 3.014 932 u.

3. Determine the binding energy for ${}^{235}_{92}\text{U}$ with a nuclear mass of 234.9934 u.

13.1 Section Review

1. K/U List the contribution of each of the following physicists to the study of the nucleus.
   (a) Ernest Rutherford
   (b) James Chadwick
2. C Why did physicists believe that a neutral particle must exist, even before the neutron was discovered?
3. K/U State the meaning of the following terms.
   (a) nucleon
   (b) atomic mass number
   (c) atomic number
4. K/U For the atom symbolized by $^{200}_{80}Hg$, state the number of
   (a) nucleons
   (b) protons
   (c) neutrons
   (d) electrons, if the atom is electrically neutral
   (e) electrons, if the atom is a doubly charged positive ion
5. K/U How did scientists know that some force other than the electromagnetic force had to exist within the nucleus?
6. C Describe the characteristics of the nuclear force.
7. C Describe the general trend of stable nuclei in relation to the proton number and neutron number.
8. C Explain the concept of binding energy.
9. C Define the term "mass defect" and explain how to determine it for a given nucleus.
10. MC The structure of the atom is often compared to a solar system, with the nucleus as the Sun and the electrons as orbiting planets. If you were going to use this analogy, which planet should you use to represent an electron, based on its comparative distance from the Sun?

Atom	Solar system
radius of nucleus ≈ 1×10^{-15} m	radius of Sun ≈ 6.96×10^{8} m
radius of typical electron orbit ≈ 1×10^{-10} m	radii of planetary orbits $r_{Mercury} \approx 6 \times 10^{10}$ m $r_{Earth} \approx 1.49 \times 10^{11}$ m $r_{Jupiter} \approx 8 \times 10^{11}$ m $r_{Pluto} \approx 6 \times 10^{12}$ m